

Answers to Section 10.1 Review Questions

(Student textbook page 648)

1. It is important to keep each half-cell separate so that an instantaneous reaction does not occur, "short-circuiting" the voltmeter.
2. The salt bridge is needed to complete the circuit so that the cell can operate. The electrolyte in the salt bridge will have negative ions flow towards the anode and positive ions flow towards the cathode. This movement of ions is responsible for completing the flow of electricity in the cell. The ends of the salt bridge are plugged with cotton or glass wool to avoid the mixing of materials from the half-cell and to stop the flow of electrolytes into the half-cells.
3. Graphic organizer must display a basic understanding of how to determine which half-reaction will be the reduction and which will be the oxidation, based on the location of each half-reaction in the standard reduction potentials chart. The answer must then also demonstrate an understanding that a spontaneous reaction in a galvanic cell will have a positive cell potential.
4. **a.** net ionic equation:
$$\text{Cu}^{2+}(\text{aq}) + \text{Ni}(\text{s}) \rightarrow \text{Ni}^{2+}(\text{aq}) + \text{Cu}(\text{s})$$
$$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Ni}(\text{s}) \quad E^{\circ}_{\text{anode}} = -0.26 \text{ V}$$
$$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu}(\text{s}) \quad E^{\circ}_{\text{cathode}} = +0.34 \text{ V}$$
$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$
$$= +0.34 \text{ V} - (-0.26 \text{ V})$$
$$= +0.60 \text{ V}$$

b. $6\text{H}_2\text{O}(\ell) + 3\text{O}_2(\text{g}) + 12\text{e}^{-} \rightarrow 12(\text{OH})^{-}(\text{aq}) \quad E^{\circ}_{\text{anode}} = +0.40 \text{ V}$
$$4\text{Au}^{3+}(\text{aq}) + 12\text{e}^{-} \rightarrow 4\text{Au}(\text{s}) \quad E^{\circ}_{\text{cathode}} = +1.50 \text{ V}$$
$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$
$$= +1.50 \text{ V} - 0.40 \text{ V}$$
$$= +1.10 \text{ V}$$

c. net ionic equation: $3\text{Cu}(\text{s}) + 2\text{Al}^{3+}(\text{aq}) \rightarrow 3\text{Cu}^{2+}(\text{aq}) + 2\text{Al}(\text{s})$
$$3\text{Cu}^{2+}(\text{aq}) + 6\text{e}^{-} \rightarrow 3\text{Cu}(\text{s}) \quad E^{\circ}_{\text{anode}} = +0.34 \text{ V}$$
$$2\text{Al}^{3+}(\text{aq}) + 6\text{e}^{-} \rightarrow 2\text{Al}(\text{s}) \quad E^{\circ}_{\text{cathode}} = -1.66 \text{ V}$$
$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$
$$= -1.66 \text{ V} - 0.34 \text{ V}$$
$$= -2.00 \text{ V}$$
5. a. and b. are spontaneous since these reactions have a positive cell potential
6. It would not be possible, as the strongest oxidizing agent is fluorine gas and the strongest reducing agent is lithium, which together would produce a cell potential of 5.91 V.
7. The cell potential describes the potential difference between two electrodes of a cell, or the amount of energy on a charge as it moves between two electrodes. The cell potential is dependent on both the anode and the cathode used. The standard reduction potential is a measure of the amount of energy for only the reduction half of cell. Since reduction cannot happen without oxidation, reduction potentials are measured against a standard reference, the hydrogen half-cell, which is set at a reduction potential of 0.00V.

8. +0.99 V

Oxidation (anode): $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$

Reduction (cathode): $\text{Pd}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pd(s)}$

$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn(s)} \quad E^\circ_{\text{anode}} = -0.76 \text{ V}$

$\text{Pd}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pd(s)} \quad E^\circ_{\text{cathode}} = X$

$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$

$+1.75 \text{ V} = X - (-0.76 \text{ V})$

$X = +0.99 \text{ V}$

9. a. Oxidation: $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$

Reduction: $\text{Ag}^+(\text{aq}) + 1\text{e}^- \rightarrow \text{Ag(s)}$

Overall cell reaction:

$\text{Zn(s)} + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag(s)}$

b. +1.56 V

$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn(s)} \quad E^\circ_{\text{anode}} = -0.76 \text{ V}$

$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag(s)} \quad E^\circ_{\text{cathode}} = +0.80$

$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$

$= +0.80 \text{ V} - (-0.76 \text{ V}) = +1.56 \text{ V}$

c. Zn(s) is the reducing agent; $\text{Ag}^+(\text{aq})$ is the oxidizing agent

d. The Ag cathode increases in mass; the Zn anode decreases in mass

10. The electrode that participates in the redox reaction will either gain mass (if it is the cathode) or lose mass (if it is the anode) as metallic atoms gain or lose electrons as the cell operates. An inert electrode does not participate in the reaction, thus will not change in mass. It simply gives the half-reaction a location to occur as electrons move in the cell. This type of electrode is used when either the oxidizing agent or the reducing agent is not a solid.

11. Each half-cell potential would have been assigned a different value, but when used in calculating the overall cell potential, there would have been no difference in the calculated value

12. Graphic organizers should include steps of how to find the oxidizing agent, reducing agent and how to calculate the cell potential from the two half-reaction potentials. Encourage students to be as creative as possible in answering this question.